

Free Energy in Reactions

↓ STP ↓

1 atm 298 K

ΔG° = standard free energy change
° - means standard

* Free energy change that will occur if reactants in their standard state turn to products in their standard state

* Can't be measured directly

* $\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$ (don't forget to convert J to kJ or kJ J)

* Use Hess's Law

* There are tables of ΔG_f° (back of the book)

* Products - reactants it is a state function

* The standard energy of formation for any element in its standard state is 0.

ΔS doesn't 0 at STP (is the only one among $\Delta S, \Delta H, \Delta T$)

* Remember! - Spontaneity does not tell anything about rate.

Free energy and Pressure

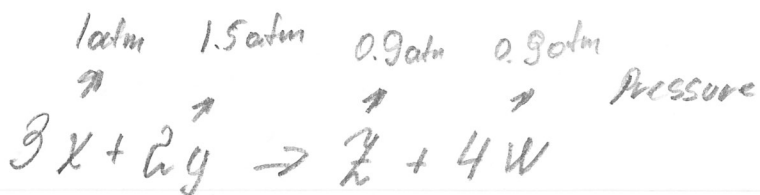
$$\Delta G = \Delta G^\circ + R \cdot T \cdot \ln(Q)$$

(use for non-constant pressure)

where Q is the reaction quotient

$$Q = \frac{P \text{ of the products}}{P \text{ of the reactants}}$$

$$Q = \frac{P_{\text{prod}_1}^{\text{coet}} \cdot P_{\text{prod}_2}^{\text{coet}}}{P_{\text{react}_1}^{\text{coet}} \cdot P_{\text{react}_2}^{\text{coet}}}$$



for this equation

$$Q = \frac{(0.9)^1 \cdot (0.9)^4}{(1)^3 \cdot (1.5)^2} = \frac{(2)^1 \cdot (W)^4}{(X)^3 \cdot (Y)^2}$$

$$\begin{array}{c}
 \Delta G = \Delta G^\circ + R \cdot T \cdot \ln(Q) \\
 \text{N/N} \quad \quad \quad \downarrow \\
 \quad \quad \quad 8.314 \text{ J/mol K}
 \end{array}
 \quad (R \cdot T \cdot \ln(Q) \text{ should be in kJ})$$

Problem:

Would the reaction be spontaneous at 25°C with the H_2 pressure of 5.0 atm and the CO pressure of 3.0 atm. (assume that CH_3OH at 1 atm because liquid)

$$\Delta G_f^\circ (CH_3OH(l)) = -166 \text{ kJ}$$

$$\Delta G_f^\circ (CO(g)) = -137 \text{ kJ}$$

$$\Delta G_f^\circ (H_2(g)) = 0 \text{ kJ}$$

ΔG_f - non-standard conditions

ΔG° - standard conditions



$$T \quad 25^\circ\text{C} \quad 25^\circ\text{C} \quad 25^\circ\text{C}$$

$$P \quad 3.0 \text{ atm} \quad 5.0 \text{ atm} \quad 1 \text{ atm}$$

$$\Delta G_f \quad -137 \text{ kJ/mol} \quad 0 \text{ kJ/mol} \quad -166 \text{ kJ/mol}$$

$$\Delta G_{rxn}^\circ = \Delta G_{prod} - \Delta G_{react} = -166 - [2(0) - (137)] = -29 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta G_{rxn} = -29 \text{ kJ/mol} + (0.008814 \text{ kJ/mol K})(298 \text{ K})(\ln(\frac{1}{3.57}))$$

$$\Delta G_{rxn} = -39.7 \text{ kJ/mol} = -40 \text{ kJ/mol}$$

ΔG is negative, therefore, the reaction is spontaneous

How Far?

- ΔG tells us about spontaneity at current conditions.
- It will go to the lowest possible free energy which may be an equilibrium.
- At equilibrium $\Delta G = 0$, $Q = K$
- $\Delta G^\circ = -R \cdot T \cdot \ln K$

ΔG°	K
$= 0$	$= 1$
< 0	> 1
> 0	< 1

$$\left[\begin{array}{l} \text{if use } Q \Rightarrow \Delta G \\ \text{if use } K \Rightarrow \Delta G^\circ \end{array} \right]$$

Relationship between ΔG° and K

$$\Delta G^\circ = -RT \ln K$$